Name ____ AP Chemistry

The Nature of Acids and Bases—Strong versus Weak

Introduction

Acids and bases are the basis for understanding many inorganic and organic processes in biochemistry, ocean chemistry, geochemistry, and many industrial processes. The Bronsted-Lowry definition of acids and bases is widely used to describe acids and bases in aqueous solutions. A Bronsted-Lowry acid is a proton donor while a Bronsted-Lowry base is a proton acceptor. Bronsted-Lowry acids and bases are related by the transfer of a proton. Water, the primary solvent in acid-base solutions, plays a key role in enabling the transfer or acceptance of protons. The involvement of water is illustrated in the reaction below for the dissolution and ionization of the strong acid HCl:

$$HCl\ (g) + H_2O\ (l) \leftrightarrow H_3O^+\ (aq) + Cl^-\ (aq)$$

Water is acting as a Bronsted-Lowry base since it is accepting a proton from HCl. Ammonia is an example of a common base. When it dissolves, it accepts a proton from water to form ammonium ion and hydroxide ion:

$$NH_3\ (g) + H_2O\ (l) \leftrightarrow NH_4^+\ (aq) + OH^-\ (aq)$$

In this reaction, water is acting as a Bronsted-Lowry acid since it donates a proton to ammonia. These two examples illustrate the importance of water as a solvent in supporting acid and base reactions.

Strong acids or bases completely dissociate in water so that no intact molecules are left. Weak acids and bases are those that partially dissociate in water, leaving a percentage of the molecules intact. The extent of dissociation of a weak acid or base is expressed mathematically by an equilibrium expression K. A reaction is said to be in equilibrium when the rate of the forward chemical reaction equals the rate of the reverse chemical reaction. Equilibrium expressions are defined to be a quotient, with the molar concentrations of the products divided by the molar concentrations of the reactants, each raised to their stoichiometric coefficient. A common weak acid is acetic acid (CH_3COOH), also known as vinegar. The acidic hydrogen is the one attached to the oxygen. This dissociation expression for acetic acid is:

$$CH_3COOH\ (aq) + H_2O\ (l) \leftrightarrow CH_3COO^-\ (aq) + H_3O^+\ (aq)$$

The resulting product of a weak acid donating its proton is referred to as the conjugate base of that acid. Thus, in this case, acetate ion, CH_3COO^-, is the conjugate base of acetic acid. The dissociation and equilibrium expression for acetic acid is given below:

$$K_a = \frac{[CH_3COO^-][H_3O^+]}{[CH_3COOH]}$$

Since the concentration of water, the solvent, does not change appreciably in the course of this reaction, it is not included in the expression; K_a is referred to as the ionization constant of a weak acid.

Note that strong acids do not have an acid ionization constant since they completely ionize in aqueous solution. As such, the term in the denominator would be zero and division by zero is undefined. The value of K_a indicates the extent of dissociation. Acid ionization constants for common weak acids range from 10^{-3} to 10^{-10}.

The dissociation and ionization of a weak base is also equilibrium process. In the balanced chemical equation of ammonia in water above, ammonium ion (NH_4^+) is the conjugate acid of the weak base, ammonia. The equilibrium quotient for this reaction is:

$$K_b = \frac{[NH_4^+][OH^-]}{[NH_3]}$$

The value of K_b indicates the extent of dissociation. Base ionization constants for common weak bases range from 10^{-3} to 10^{-10}.

The concentration of H_3O^+ and OH^- in solution can vary over many orders of magnitude from 10 M to 10^{-14} M. In 1909, Sorensen proposed the use of a logarithmic scale that we commonly use today. To denote the hydrogen ion concentration, pH is used:

$$pH = - \log [H_3O^+]$$

Similarly, to denote the hydroxide ion concentration, pOH is used:

$$pOH = - \log [OH^-]$$

This logarithmic "shorthand" of the molar concentration is actually a "p operator" in math terminology. The letter p comes from the German word *potenz* which means power or exponent of (power of 10 in this case). The "p" notation always implies that the result is derived by taking the negative logarithm (base 10).

Previously, the special nature of water to act as both an acid and a base was noted. In fact, water itself undergoes an acid-base reaction:

$$2\ H_2O\ (l) \leftrightarrow H_3O^+\ (aq) + OH^-\ (aq)$$

The equilibrium constant for this reaction is:

$$K_w = [H_3O^+][OH^-]$$

Since the value of K_w at 25ºC is 10^{-14}, a solution with a pH of 7.00 is considered a neutral solution. Solutions with a pH less than 7 are acidic while solutions with pH greater than 7 are basic.

The measurement of pH is mainly done by two methods:

- pH indicators: A pH indicator is an organic weak acid whose conjugate base is a much different color than that of the acid. The weak acid/base can be added to the solution or impregnated on a piece of paper. Comparison of the color of the solution (or the wetted piece of paper impregnated with the indicator) to a color scale for the indicator enables estimation of the solution's pH. A generalized reaction of a weak acid indicator is

  $$HIn\ (aq) \leftrightarrow H^+\ (aq) + In^-\ (aq)$$

  where HIn denotes its acidic form and In^- denotes the conjugate base form. Usually, the ratio of weak acid [HIn] to weak base [In^-] must be at least 1:10 or 10:1 for our eye to be able to detect the color change (1 part in 10). Outside this range, it is difficult to sense a change in color or hue with your eyes.

- pH meter: A pH meter measures the voltage developed across a thin glass membrane due to the difference in H^+ ion concentration inside the bulb versus outside the bulb (relative to a reference voltage/electrode). This voltage is converted to a meaningful pH scale and displayed on the pH meter. This small voltage measurement is very sensitive to temperature and has a response time of several seconds. The pH meter must be calibrated with solutions of known pH for accurate measure of solution pH.

Prelab Questions

1) Calculate the pH for the acids and pOH for the bases. For strong acids and bases assume complete dissociation.

Strong Acid	**pH**	**Strong Base**	**pOH**
1 M HCl		1.4 M KOH	
0.01 M HCl		4 M KOH	
0.001 M $HClO_4$		0.01 M NaOH	
1.5×10^{-4} M HNO_3		0.001 M NaOH	
4.6×10^{-6} M $HClO_4$		1×10^{-12} M NaOH	

2) Describe the two different methods of measuring pH.
3) Name two household items that are basic and two items that are acidic.

Procedure

- Part A: Common Acids and Bases
  1) Obtain a well plate and place a small portion of each household item in separate wells.

2) Measure the pH of each item with a strip of pH paper and then with one drop of universal indicator. Record the color and pH in Data Table 1.
3) Record each item in Data Table 1 along with the main component of each item.

- Part B: Acid/Base Strength
  1) Go to the following website: http://group.chem.iastate.edu/Greenbowe/sections/projectfolder/flashfiles/acidbasepH/ph_meter.html
  2) Choose “Acid” in the lower left corner. Choose “H_2SO_4” from the list of acids and set the molarity to “10^{-2}” as the choice and 1 in the box.
  3) Click “Insert Probes” on the pH meter and record the pH in Data Table 2.
  4) Click “Remove Probes.” Change molarity to “10^{-3}” and click “Insert Probes.”
  5) Record the pH and click “Remove Probes.”
  6) Change molarity to “10^{-3}” and click “Insert Probes.”
  7) Record the pH and click “Remove Probes.”
  8) Continue to measure the pH of each concentration of $HC_2H_3O_2$.
  9) Choose “Base” from the lower left corner and record the pH for both bases at all concentrations in Data Table 2.
  10) Choose “Salt I” and measure pH of 1.0 x 10^{-2} M of each salt. Record the pH in Data Table 3.
  11) Repeat step #10 with “Salt II” and “Salt III.”

Data Tables

Data Table 1

Product/Compound	Active Ingredient	pH Paper	pH Indicator

Data Table 2

Compounds	**1.0 x 10^{-2} M**			**1.0 x 10^{-3} M**			**1.0 x 10^{-6} M**		
Acids	*pH*	*[H^+]*	*% Dissociation*	*pH*	*[H^+]*	*% Dissociation*	*pH*	*[H^+]*	*% Dissociation*
H_2SO_4			N/A			N/A			N/A
$HC_2H_3O_2$									
Bases	*pOH*	*[OH^-]*	*% Dissociation*	*pOH*	*[OH^-]*	*% Dissociation*	*pOH*	*[OH^-]*	*% Dissociation*
NaOH			N/A			N/A			N/A
NH_3									

Data Table 3

Compound	**pH**	**Acidic or Basic?**	**Compound**	**pH**	**Acidic or Basic?**
NaCl			K_2SO_4		
NH_4NO_3			KCN		
$NaC_2H_3O_2$			KNO_3		
NH_4Cl			KF		
Na_2SO_4			K_2CO_3		
NaCN			Na_3PO_4		
$NaNO_3$			$NaHCO_3$		
NaF			$NaHSO_4$		
$(NH_4)_2SO_4$			$NH_4C_2H_3O_2$		
KCl			NH_4F		
$KC_2H_3O_2$					

Postlab Questions

1) Calculate the percent dissociation for acetic acid and ammonia using the following formulas:

$$\%\,Dissociation = \frac{[H^+]}{[weak\ acid]} x\ 100\ OR\ \frac{[OH^-]}{[weak\ base]} x100$$

2) Discuss the pH of common household items. Which item was the strongest acid? Strongest base? What ingredients cause the acidity or basicity?
3) Discuss what happens to pH/pOH, $[H^+]/[OH^-]$, and percent dissociation for acetic acid and ammonia as the compound is diluted. Did the percent dissociation increase or decrease as the solutions are diluted?
4) What trends can be observed in the pH of the salts? What cation or anion is more acidic? More basic?